STOICHIOMETRY

Chapter 1

Teaching Periods	Assessment	Weightage
10	1	8

Students will be able to:

- **Describe** mole and Avogadro's number with examples.
- **Determine** Avogadro's number and describe the relationship between moles and Avogadro's number.
- **Define** rounding off data, exponential notation and their practical applications in solving numericals.
- **Perform** stoichiometric calculations with balanced equations using moles, representative particles, masses and volumes of gases (at STP).
- **Identify** the limiting reactant in a reaction.
- **Know** the limiting reactant in a reaction. Calculate the maximum amount of product(s) produced and the amount of any unreacted excess reactant.
- **Calculate** theoretical yield, actual yield and percent yield from the given information.
- **Calculate** theoretical yield and percent yield by using balance equation.

INTRODUCTION

The spirit of chemistry is recognized in its practical nature. Just imagine how useful it could be to determine the amount of products formed when some known quantities of reactants undergo a chemical change. Think about an automotive engineer studying flue gases: what amount of exhaust gases will be produced due to the combustion of gasoline in an internal combustion engine? Or suppose you are in your kitchen to make 12 cups of delightful coffee with adequate quantity of sugar and milk. You need to calculate the proportion of each ingredient to be added. You won't be able to conceive the answers to these questions without the knowledge of **stoichiometry**.

Stoichiometry (Greek Stoicheion; element, and metron; measurement) deals with the study of quantitative relationship between reactants and products in a chemical reaction by using balanced chemical equation. You know that oxygen gas and hydrogen gas chemically combines to form water but without the knowledge of stoichiometric amounts of hydrogen (H_2) and oxygen (O_2), you cannot estimate how much water (H_2O) will be formed.

Conclusively, stoichiometry tells us what amount of each reacting species we require to consume completely into desired amount of product(s). Although stoichiometry is an important tool for solving diverse problems in laboratory, chemical industries, engineering, food manufacturing, pharmaceuticals and many other fields of science but it is based on the assumption that reactant molecules are completely converted into product. In fact, many chemical reactions are reversible to some extent, further, in some chemical processes side reactions also occur.

1.1 MOLE AND AVOGADRO'S NUMBER

In routine work, we generally measure things by weighing or by counting with the option based on our comfort. It is more convenient to buy gloves by pairs (two gloves), bananas by dozens (one dozen is equal to 12 bananas) tea sachets by gross (one gross is equal to 144 sachets) and paper by reams (one ream is equal to 500 sheets) but purchasing rice from a grocery shop is convenient by weighing instead of counting.

Similarly, if we prepare a solution or perform a reaction in chemistry laboratory or even in process industry we deal with the enormous numbers of ions, molecules or formula units that mix with one another in specific ratio. Since atom or molecule is so tiny, how can it be possible to count them? To solve this difficulty chemists have devised a unit called as **mole** as a convenient way to count the number particles in chemical substance by weighing them.

Mole (Latin: heap or pile) is the SI unit use for measuring the amount of substance of specific number of particles. **"A mole is defined as gram atomic mass or gram molecular mass or gram formula mass of any substance (atoms, molecules, ions) which contains 6.02×10^{23} particles".** Mole represents the number of chemical entities in a fixed mass.

Conversely, we can say that one mole of any of the substances always contains 6.02×10^{23} particles and known as Avogadro's number (N_A) in the honor of Italian physicist Amedeo Avogadro. **"The number of particles present in one mole of any substance is called Avogadro's number (N_A) and its numerical value is 6.02×10^{23}"**.

One mole of any substance always contains 6.02×10^{23} particles and no matter what the chemical nature of a substance is? Thus, one mole of carbon and one mole of magnesium contain same number of atoms but one mole of magnesium has a mass twice (24g) as that of one mole of carbon (12g).

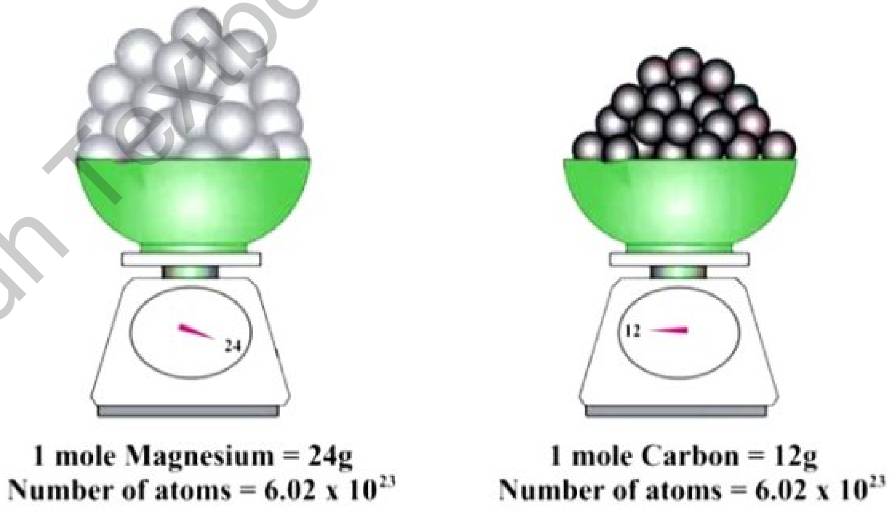

Fig. 1.1: one mole quantities of different substances contains equal number of particles

While using the term mole for ionic compounds, we consider the number of formula units of that compound. For example one mole of $MgCl_2$ is a quantity which contains 6.02×10^{23} $MgCl_2$ units. However, in each $MgCl_2$ unit there is one Mg^{+2} ion and two Cl^{-} ions. This indicates that one mole of $MgCl_2$ contains 6.02×10^{23} Mg^{+2} ion and 12.04×10^{23} Cl^{-} ions.

$$MgCl_2 \rightarrow Mg^{+2} + 2Cl^{-}$$

Do You Know?

Avogadro's number (6.02×10^{23}) is such a huge value that if the number of atoms of an element is counted by a device at the rate of one million per second, it would take four billion years similarly if there is such number of basketballs it would make entirely a new planet like earth.

Comparing a dozen of bananas and a mole of carbon, it is important to note that bananas vary in mass but all carbon atoms have equal mass. Thus a fruit seller cannot sale one dozen bananas by weighing them however a chemist can deal with 1 mole carbon easily by weighing 12g of carbon. **"The mass in grams of one mole of any pure substance is known as molar mass"**. The unit for molar mass is grams per mole (g/mol).

The basic difference between the mass of one atom and the mass of 1 mole is that the atomic mass of one atom of an element is specified by a.m.u and the mass of one mole of an element is expressed in grams. Thus atomic mass of Fe is 55.85 a.m.u but its molar mass is 55.85 gram. A similar relationship holds for compounds for instance molecular mass of propane (C_3H_8) is 44 a.m.u but its molar mass is taken as 44 grams.

The volume of the given quantity of gas may be different at different temperatures and pressures. Chemists use standard conditions of temperature and pressure for taking the volume of gas. **"The volume of one mole of a gas at standard temperature (273K) and pressure (1 atm) is referred as molar volume"**. Molar volume of all ideal gases at STP is $22.4dm^3$ and can be determined by dividing molar mass with mass density. (molar volume = molar mass /density).

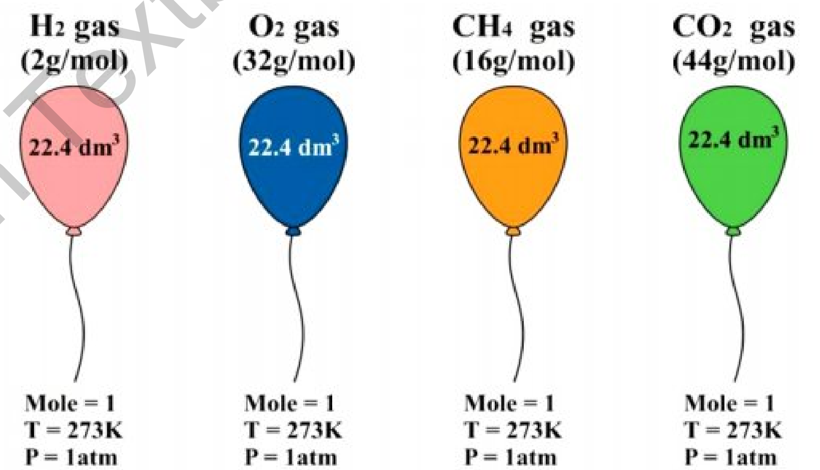

Fig 1.2: Molar volume of different gases

Table 1.1	Relationship between Mole, Molar mass and Avogadro's Number		
Name of substance	**Symbol/ Formula**	**Molar Mass (g/mol)**	**Kind and number of particles in one mole**
Sodium	Na	23.0	6.02×10^{23} Na atoms
Water	H_2O	18.0	6.02×10^{23} H_2O molecules 6.02×10^{23} O atoms $2(6.02 \times 10^{23})$ H atoms
Common Salt	NaCl	58.5	6.02×10^{23} NaCl Formula units 6.02×10^{23} Na^+ ions 6.02×10^{23} Cl^- ions
Calcium Chloride	$CaCl_2$	111	6.02×10^{23} $CaCl_2$ Formula units 6.02×10^{23} Ca^{2+} ions $2(6.02 \times 10^{23})$ Cl^- ions
Ferric Chloride	$FeCl_3$	162.35	6.02×10^{23} $FeCl_3$ Formula units 6.02×10^{23} Fe^{3+} ions $3(6.02 \times 10^{23})$ Cl^- ions

Inter conversion of Mole and Mass

While working on stoichiometric problems, we often need to interconvert mole and mass of reactants and products. For this purpose molar mass is used as conversion factor. To convert given mass of substance into mole, we divide it by molar mass. On the other hand if moles are needed to be converted into mass, we simply multiply it with molar mass. Conclusively, the following formula may be used for the inter conversion of mole and mass.

$$\text{No. of moles} = \frac{\text{Given mass (g)}}{\text{Molar mass of substance}}$$

Example 1.1

Calculate the number of moles in 25.5g of sodium metal.

Data:

Given mass of sodium metal =25.5 g

No. of moles of sodium metal=?

Solution:

Since the molar mass of sodium is 23g/mol, we use this molar mass as conversion factor to determine the number of moles of sodium.

$$\text{Moles of Na} = \frac{\text{Given mass of Na}}{\text{Molar mass of Na}} = \frac{25.5\text{g}}{23\text{g/mol}} = 1.11 \text{ moles}$$

Example 1.2

Calculate the mass of 3.25 moles of water (H_2O).

Data:

Given No. of moles of water = 3.25

Mass of water (H_2O) = ?

Solution:

To convert mass of water into moles, we use the following relation

$$\text{Moles of water} = \frac{\text{mass of water}}{\text{molar mass of water}}$$

Mass of water = moles of water × molar mass of water

Mass of H_2O = 3.25 × 18 = 58.5g

Inter conversion of Mole and Number of Particles

If we need to convert moles into number of particles or vice versa, Avogadro's number is used as converting factor. This can be expressed in the following way.

$$\text{No. of moles} = \frac{\text{No. of particles of the substance}}{\text{Avogadro's Number}}$$

Example 1.3

Calculate the number of molecules in 610g of Benzoic acid ($C_7H_6O_2$)

Data:

No. of molecules of Benzoic acid ($C_7H_6O_2$) = ?

Given Mass of Benzoic acid ($C_7H_6O_2$) = 610g

Solution:

Since molar mass of benzoic acid is 122 g/mol

122 g benzoic acid ($C_7H_6O_2$) contains = 6.02×10^{23} molecules of benzoic acid

1g .... $= \dfrac{6.02\times10^{23}}{122}$

610g .... $= \dfrac{6.02\times10^{23}}{122}\times 610$

$= 30.1 \times 10^{23}$ molecules of benzoic acid

Example 1.4

Calculate the mass of 4.39 × 10^{24} atoms of Gold(Au), molar mass of gold is 197 g/mol

Data:

Mass of Gold (Au) atoms =?

Given No. of atoms of Gold (Au) = 4.39 × 10^{24} atoms

Given Molar mass of Gold (Au) = 197 g/mol

Solution:

6.02 × 10^{23} atoms of Gold (Au) = 197 g of Gold

1 atom ……………………. $= \dfrac{197}{6.02\times10^{23}}$ g of Gold

4.39 × 10^{22} ………………..$= \dfrac{197\times4.39\times10^{22}}{6.02\times10^{23}}$ =14.365 g of Gold

Example 1.5

Calculate the number of moles in 2.35 × 10^{25} atoms of Aluminum (Al).

Data:

No. of moles of Aluminum (Al) atoms = ?

Given No. of atoms of Aluminum (Al) = 2.35 × 10^{25} atoms

Solution:

Since one mole of any substance contains = 6.02 × 10^{23} particles.

Hence

6.02 × 10^{23} atoms of Al = 1 mole of Al

1 atom ………………….. $= \dfrac{1}{6.02\times10^{23}}$

2.35 × 10^{25} …………….. $= \dfrac{1}{6.02\times10^{23}}\times2.35\times10^{25} = 39$ moles of Al

Self Assessment

Graphite is one of the two crystalline forms of carbon which is a constituent component of lead pencils. How many atoms of carbon are there in 360 g of graphite? Also find the number of moles of carbon.

Inter conversion of Mole and Molar Volume

In certain stoichiometric calculations where gas evolves in the reaction, we are asked to determine the volume of gas released. In this situation we use molar volume as conversion factor.

$$\text{No. of moles} = \frac{\text{Volume of gas occupied at STP}}{\text{Molar Volume}}$$

Example 1.6

What volume of oxygen gas (O_2) occupied by 1.5 moles at STP.

Data:

Volume of oxygen gas (O_2) = ?

Given No. of Moles of oxygen gas (O_2) at STP = 1.5 moles

Solution:

Since the volume of 1 mole of O_2 at STP is 22.4 dm^3

$$\text{No. of moles } O_2 = \frac{\text{Volume of } O_2}{22.4}$$

Volume of O_2 (STP) = 1.5 × 22.4 = 33.6 dm^3

Calculations based on mass–mass relationship

In these types of problems, we determine the unknown mass of reactant or product from the given mass of substance in a chemical reaction with the help of balanced chemical equation.

Example 1.7

Calculate mass of carbon dioxide (CO_2) that can be obtained by complete thermal decomposition of 50g Lime stone ($CaCO_3$).

$$CaCO_{3(s)} \xrightarrow{\text{Heat}} CaO_{(s)} + CO_{2(g)}$$

Data:

Mass of carbon dioxide (CO_2) =?

Given mass of Lime stone ($CaCO_3$) = 50g

Solution:

No. of moles in 50g $CaCO_3$ $= \frac{50}{100} = 0.5$ moles

1 mole of $CaCO_3$ gives ……. = 1 mole of CO_2

0.5 moles ………………… = 0.5 moles of CO_2

Now, mass of CO_2 ……….. = moles of CO_2 × molar mass of CO_2

Mass of CO_2 = 0.5 × 44 = 22g

Calculation based on mass–volume relationship

Many chemical reactions are carried out with the production of gas from a solid substance. In this case the measurement of volume is more appropriate than mass. Stoichiometry of this type is known as mass-volume or volume-mass relationship.

Example 1.8

Mass of 49g of solid potassium chlorate ($KClO_3$) on heating decomposes completely to potassium chloride (KCl) with the liberation of oxygen gas (O_2).

$$2KClO_{3(s)} \xrightarrow{\text{Heat}} 2KCl_{(s)} + 3O_{2(g)}$$

Determine volume of oxygen gas (O_2) liberated at STP

Data:

Given mass of potassium chlorate ($KClO_3$) = 49g
Volume of oxygen gas (O_2) =?

Solution:

No. of moles of $KClO_3 = \frac{49}{122.5} = 0.4$ moles

2 moles of $KClO_3$ gives = 3 moles of O_2

1 .... $= \frac{3}{2}$ moles of O_2

0.4 ....$= \frac{3}{2} \times 0.4 = 0.6$ moles of O_2

Now the volume of O_2 at STP can be determined by using Avogadro's concept.

Volume of 1 mole of a gas at STP = 22.4dm^3
Volume of 0.6 mole of a oxygen gas (O_2) at STP = 22.4 x 0.6 dm^3
Volume of O_2 at STP = 13.44 dm^3

Self Assessment

On heating solid ammonium nitrate (NH_4NO_3), it decomposes to produce nitrous oxide (N_2O) and water

$$NH_4NO_{3\,(s)} \rightarrow N_2O_{\,(g)} + 2H_2O_{\,(l)}$$

If 200g of ammonium nitrate is completely consumed in the reaction calculate the:
(i) Mass of water formed (ii) Volume of N_2O gas liberated at STP

Calculations based on volume–volume relationship

In these types of problems, we determine the unknown volume of a reactant or product from the given volume of substance in a chemical reaction with the help of balanced chemical equation.

Example 1.9

Calculate the volume of carbon dioxide at STP that can be produced by the complete burning of 50 dm^3 of butane gas (C_4H_{10}) in the excess supply of oxygen gas (O_2).

$$2C_4H_{10(g)} + 13O_{2(g)} \xrightarrow{\text{Heat}} 8CO_{2\,(g)} + 10H_2O_{(g)}$$

Data:

Volume of carbon dioxide (CO_2) at STP =?

Given volume of Butane (C_4H_{10}) = 50 dm^3

Solution:

According to balanced chemical equation:

2 dm^3 of butane produces = 8 dm^3 of carbon dioxide

1 dm^3…………………… $= \frac{8}{2}$

50 dm^3 ………………… $= \frac{8}{2} \times 50 = 200$ dm^3 of CO_2

1.2 ROUNDING OFF DATA

In chemical calculations our result is often consist of too many figures and for convenience we often round off the answer into proper numbers by dropping last digit(s). **"To reduce a number upto desired significant figures and adjust the last reported digit is known as rounding off data"**. The right most digits is generally considered as uncertain therefore, we conveniently drop it and round off the figure into smaller numbers to ensures the maximum removal of errors from the final result.

Generally, numbers are rounded to the nearest ten, hundred, thousand, million and so on.

Do You Know?

The digits in a number which show reliability in measurement are known as significant figures.

- All non-zero digits are significant figures.
- Zeros lying in between non-zero digits are significant figures.
- Zeros locating right after the decimal point in number less than one are not significant.
- Final zeros to the right of the decimal point are significant.
- Zeros that locate before the decimal point in number less then one are not significant.

Rules for rounding off data:

(i) If digit to be dropped is greater than 5, then add 1 to the digit to be retained.

For example: 5.768 is rounded up to 5.77 if three significant figures are needed to be retained.

(ii) If digit to be dropped is less than 5, then simply drop it without changing preceding number.

For example: if 5.734 is rounded up to three significant figures, we get 5.73

(iii) If digit to be dropped is exactly 5, there are two conditions:

* If the digit to be retained is even, then just drop the 5.

For example: when 7.865 is rounded up to three figures we get 7.86

* If the digit to be retained is odd, then add 1 to it

For example: 23.35 is rounded to 23.4

1.3 EXPONENTIAL NOTATIONS

While dealing scientific work we often face difficulties in the calculations of very small or very large numbers. Often we found wrong results due to the mistake in writing digits of too small and too large numbers. These numbers are much more conveniently expressed as multiple of 10.

For example, the mass of electrons is 0.0000000000,0000000000,0000000 911g which is written conveniently as 9.11×10^{-28}g. Similarly value of Avogadro's number is 602,300,000,000,000,000,000,000 which is written in exponential notation as 6.023×10^{23}.

Exponential notation not only helps in the simplification of calculations but also save space and time. Due to convenience, it is frequently used in different fields like engineering, physics, geology, astronomy, etc.

Exponential notation is represented as the product of two numbers $N \times 10^x$.

$N \times 10^x$

Coefficient **Base** **Exponent**

Where N (coefficient) may be the number between 1 to 9.999…

x (exponent) is an integer raised to base 10. Exponent may be positive or negative.

If the number to be expressed in exponential notation is greater than 1, the exponent is positive integer (x > 1) but if it is less than 1 (x < 1), the exponent is negative integer.

Inter Converting Standard and scientific notation

In order to tackle various mathematical calculations in chemistry you must aware how to convert standard notation into exponential notation or vice versa.

Consider the following two rules

(i) If we need to convert a numerical value from standard notation to exponential notation, decimal point should move to the left if the value is greater than 10 and move to right if the value is in between 0 to 1.

For example:

- 4600,000 change to 4.6×10^6 (decimal point moves six place to the left)
- 0.00038 change to 3.8×10^{-4} (decimal point moves four place to the right)

(ii) If we want to change an exponential value into standard notation, the decimal point should move to the right for the positive exponent and move to the left for negative exponent.

For example:

- 7.53×10^4 can convert into 75300 (decimal point moves four place to the right)
- 48.7×10^{-5} can convert into 0.000487 (decimal point moves five place to the left)

Applications of exponential notations

(i) Addition and subtraction

Before addition or subtraction, convert all numbers into the same exponents of 10. Then add or subtract the digit terms (coefficients).

Example 1.10

Add 1.31×10^3 and 3.15×10^2 by using the rule of exponential notation.

Solution:

The value 3.15×10^2 is initially converted into 0.315×10^3 by placing decimal point to the left. Then add the coefficients of both values.

$$\begin{array}{r} 1.31 \times 10^3 \\ \underline{+\,0.315 \times 10^3} \\ 1.625 \times 10^3 \end{array}$$

(ii) Multiplication

Multiply all digit terms (coefficients) and add all exponents algebraically. The final result may be adjusted by placing the decimal to the left or right.

Example 1.11

Multiply 7.0×10^{12} and 2.0×10^{-3} by using the rule of exponential notation.

Solution:

Coefficients 7.0 and 2.0 will be multiplied while exponents 10^{12} and 10^{-3} will be algebraically added.

$$= (7.0)\,(2.0) \times 10^{12-3}$$
$$= 14 \times 10^9$$
$$= 1.4 \times 10^{10}$$

(iii) Division

Divide all digit terms (coefficients) and subtract all exponents algebraically. The final result may be adjusted by placing the decimal to the left or right.

Example 1.12

Divide 6.60×10^8 with 3.20×10^3 by using the rule of exponential notation.

Solution:

Coefficient 6.60 and 3.20 will be divided while exponents 10^8 and 10^3 will be algebraically subtracted.

$$= \frac{6.60 \times 10^8}{3.20 \times 10^3}$$
$$= \frac{6.60 \times 10^{8-3}}{3.20}$$
$$= 2.60 \times 10^5$$

(iv) Powers

Multiply the digit terms as well as exponent with a number which indicates the power. The final result may be adjusted by placing the decimal to the left or right.

Example 1.13

Simplify $(3.25 \times 10^4)^2$ by using rules of exponential notation.

Solution:

Here digit term is 3.25 and exponent term is 10^4. Both are multiplied by whole power of the figure to get the answer.

$= (3.25)^2 \times 10^{4x2}$

$= 10.56 \times 10^8$

$= 1.056 \times 10^9$

(v) Roots

Adjust the exponent by shifting the decimal to the right or left in digit term in such a way that exponent must exactly be divided by the root. Then simplify the root of digit term and divide the exponent by a desired root.

Example 1.14

Simplify $\sqrt{2.5\times10^7}$ by using rules of exponential notation.

Solution:

Digit term is 2.5. It is adjusted to 25 by placing the decimal to the right.

$= \sqrt{25\times10^6}$

Now root of both digit term and exponent can be taken to get the answer.

$= 5.0 \times 10^3$

1.4 LIMITING REACTANT AND ITS CALCULATIONS

We know that a sandwich is made of two slices of bread and one shami. Suppose we have 20 slices of bread and 8 shamies. From this quantity we will be able to make only 8 sandwiches. Thus, 4 slices of bread will be left over. The available quantity of shamies limit the number of sandwiches.

The similar situation is occurred in many irreversible chemical reactions where one reactant is often completely used while some amount of other reactant remains unreacted. **"The reactant which is entirely consumed first during chemical reaction is called Limiting reactant or Limiting reagent".**

The reactant that is not completely consumed is often referred as excess reactant. Mathematically, limiting reactant is that which gives least number of moles of product in the chemical reaction because once, one of the reactants is consumed completely, the reaction stops. Consider the formation of water from hydrogen gas (H_2) and oxygen gas (O_2) in the following balance equation.

$$2H_{2\,(g)} + O_{2\,(g)} \longrightarrow 2H_2O_{\,(l)}$$

If we start the reaction by taking 2 moles of hydrogen and 2 mole of oxygen it will be noted that hydrogen is consumed earlier and stops the reaction while O_2 left behind. Thus, H_2 is considered as "Limiting reactant" and O_2 is identified as "excess reactant".

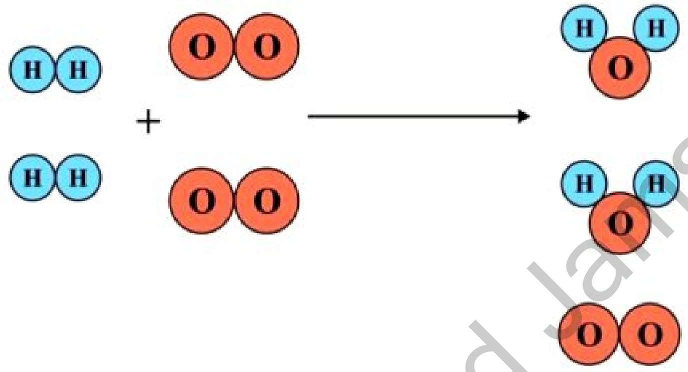

Fig 1.3: Reaction of two moles of H_2 and two moles of O_2 to give H_2O shows that hydrogen is limiting reactant while O_2 is excess reactant.

To find out a limiting reactant in a chemical reaction we must focus on following four steps.
(i) Write a balanced chemical equation of the given chemical process.
(ii) Determine the number of moles of reactants from their given amount.
(iii) Find out the number of moles of product with the help of a balanced chemical equation.
(iv) Identify the reactant which produces the least moles of product as limiting reactant.

Example 1.15

Combustion of Ethene in air to form CO_2 and H_2O is given in the following equation.

$$C_2H_{4(g)} + 3O_{2(g)} \longrightarrow 2CO_{2(g)} + 2H_2O_{(l)}$$

If a mixture containing 2.8g C_2H_4 and 6.4g O_2 is allowed to ignite, identify the Limiting reactant and determine the mass of CO_2 gas will be formed.

Data:

Mass of Ethene (C_2H_4) = 2.8g
Mass of Oxygen gas (O_2) = 6.4g
Mass of carbon dioxide (CO_2) =?
Limiting reactant =?

Solution:

To solve this problem you first convert the given masses of both reactants into their moles.

$$\text{Mole of } C_2H_4 = \frac{2.8}{28} = 0.1$$

$$\text{Mole of } O_2 = \frac{6.4}{32} = 0.2$$

To find out whether C_2H_4 or O_2 consumed earlier, we should go for the following calculations.

Mole comparison of C_2H_4 and CO_2

According to balanced chemical equation

1 mole of C_2H_4 gives = 2 moles of CO_2

0.1 mole ……………. = 0.2 moles of CO_2

Mole comparison of O_2 and CO_2

According to balance chemical equation

3 moles of O_2 gives = 2 moles of CO_2

1 ………………… $= \frac{2}{3}$ moles of CO_2

0.2 ………………. $= \frac{2}{3} \times 0.2 = 0.133$ moles of CO_2

Since the number of moles of CO_2 produced by O_2 is less, therefore O_2 is Limiting reactant.

Now: Amount of CO_2 produce will be calculated as

Amount of CO_2 = moles of CO_2 x molar mass of CO_2

Amount of $CO_2 = 0.133 \times 44 = 5.852$g

Example 1.16

When Aluminum is heated with Nitrogen gas at 700 °C, it gives Aluminum nitride.

$$2Al_{(s)} + N_{2(g)} \longrightarrow 2AlN_{(s)}$$

If 67.5 g of Aluminum and 140g of Nitrogen gas are allowed to react, find out the:

(a) Limiting reactant?

(b) Mass of Aluminum nitride (AlN) produced?

(c) Mass of excess reactant?

(Atomic mass of Al is 27 a.m.u and N is 14 a.m.u)

Solution:

We first convert the given amount of reactants into moles.

$$Al = \frac{62.5}{27} = 2.5 \text{ moles}$$

$$N_2 = \frac{140}{28} = 5 \text{ moles}$$

Mole comparison of Al and AlN

According to balanced chemical equation

2 moles of Al gives = 2 moles of AlN

1 ………………… $= \frac{2}{2}$

2.5 ……………… $= \frac{2}{2} \times 2.5 = 2.5$ moles

Mole comparison of N_2 and AlN

1 mole of N_2 gives = 2 moles of AlN

5 moles .... = 2 × 5 = 10 moles

Since number of moles of AlN produced by Aluminum is less, therefore Limiting reactant is Aluminum

Amount of AlN is calculated by multiplying its moles with molar mass

Mass of AlN = 2.5 × 41 = 102.5g

Now, mass of excess reactant is determined as

2 moles of Al combine with = 1 mole of N_2

1 .... $= \frac{1}{2}$ mole of N_2

2.5 .... $= \frac{1}{2} \times 2.5 = 1.25$ moles of N_2

Excess moles of N_2 = 5 – 1.25 = 3.75moles

Excess amount of N_2 = 3.75 × 28 = 105g

Self Assessment

Hydrogen gas is commercially prepared by steam methane process.

$$CH_4 + H_2O \xrightarrow{\text{Heat}} CO + 3H_2$$

If a mixture of 28.8g methane and 14.4g steam is heated in a furnace at elevated temperature to liberate carbon monoxide and hydrogen gases. Determine the Limiting reactant and the mass of hydrogen gas produced.

1.5 THEORITICAL YIELD, PRACTICAL YIELD AND PERCENT YIELD

After getting a proper knowledge about the concept of stoichiometry and Limiting reactant, we now focus on another aspect of calculation in chemical process, this is known as reaction yield. There are three types of yields:

(i) Theoretical yield
(ii) Actual yield
(iii) Percentage yield

The maximum amount of product obtained by a balanced chemical reaction by using its Limiting reactant is known as **Theoretical Yield**.

No matter how much you have expertise in chemical reactions or you use highly efficient techniques; you will lose some amount of product during the course of reaction in laboratory or chemical plant. Thus what we actually obtain is less than what we calculate (theoretical yield). The actual amount of product which is formed in experiment is called **Practical or Actual yield**. This difference between theoretical and practical yield is due to various causes.

(i) Either some amount of reactant may not react
(ii) The reactants may form any side product.

(iii) There may occur reversible reaction.

(iv) There may loss certain amount of product due to physical processes like distillation, filtration, crystallization and washing etc.

The efficiency of a chemical process is judged by calculating the ratio of practical yield to theoretical yield. This ratio is known as **percent yield**.

$$\text{Percentage yield} = \frac{\text{Practical Yield}}{\text{Theoretical Yield}} \times 100$$

Example 1.17

The reaction of calcium carbonate ($CaCO_3$) with hydrochloric acid is given as

$$CaCO_{3(s)} + 2HCl_{(aq)} \longrightarrow CaCl_{2(s)} + CO_{2(g)} + H_2O_{(l)}$$

If during an experiment 50 g of $CaCO_3$ is reacted with excess of hydrochloric acid, 14.52 g of CO_2 gas is liberated, calculate the theoretical and percentage yield of CO_2 gas.

Solution:

We first determine the number of moles of $CaCO_3$

$$\text{Moles of } CaCO_3 = \frac{50}{100} = 0.5 \text{ moles}$$

According to balanced chemical equation

1 mole of $CaCO_3$ gives = 1 mole of CO_2

0.5…………………… = 0.5 moles of CO_2

Theoretical yield of CO_2 is now determined by multiplying its number of moles with molar mass

Theoretical yield of CO_2 = 0.5 × 44 = 22g

Percentage yield can be determined by using the formula.

$$\text{Percentage yield} = \frac{\text{Practical Yield}}{\text{Theoretical Yield}} \times 100$$

$$\text{Percentage yield of } CO_2 = \frac{14.52}{22} \times 100 = 66\%$$

Self Assessment

Under high pressure Magnesium (Mg) reacts with oxygen (O_2) to form Magnesium oxide (MgO).

$$2Mg + O_2 \longrightarrow 2MgO$$

If 4 grams of Mg reacts with excess of O_2 to produce 4.24 g of MgO. Calculate the percentage yield of MgO.

Society, Technology and Science

Limiting Reactants in the preparation of bakery items

Preparing bakery items is a suitable analogy to comprehend the fundamental concept of stoichiometry.

For example, a recipe for making sixteen pancakes required for two cups of pancake powder, one and a half cup of milk, and two eggs. The equation expressing the preparation of sixteen pancakes is

2 cup pancake powder + 1.5 cup milk + 2 eggs → 16 pancakes

If four dozen pancakes are needed for family breakfast when large number of family members gets together, the ingredient quantity must be increased proportionally according to the amounts stated in the recipe. For example, the number of eggs required to make 48 pancakes is (2 eggs / 16 pancakes) × (48 pancakes) = 6 eggs

Activity

This activity is aimed to calculate the number of molecules of a substance in common use. You may take one or two teaspoons of sugar ($C_{12}H_{22}O_{11}$) when you have a cup of tea. Suppose you take 2 teaspoons of sugar, can you estimate the number of sugar molecules that you drink along with the tea?

In order to find out the number of molecules of sugar, first of all weigh the sugar on a digital balance of your kitchen, then find out the number of moles of sugar by dividing the weighed mass by the molar mass of sugar (342 g/mol). Now you can estimate the number of sugar molecules by multiplying the number of moles with Avogadro's number (6.02×10^{23}).

SUMMARY with Key Terms

- **Stoichiometry** is the study of quantitative relationship between reactants and products in a chemical reaction by using balanced chemical equation.
- **Mole** is the gram atomic mass or gram molecular mass or gram formula mass of any substance (atoms, molecules, ions) which contains 6.02 x 10^{23} particles.
- **Avogadro's Number** is the number of particles (atoms, ions, molecules) present in one mole of any substance. It is denoted by N_A and its value is 6.02 x 10^{23}.
- **Molar Volume** is the volume occupied by one mole of any gas at 273K temperature and 1 atmospheric pressure. Molar volume of all ideal gases at STP is 22.4dm^3 and can be determined by dividing molar mass with density.
- **Molar Mass** is the mass of one mole of any substance; it is measured in g/mol.
- **Rounding off data** is a method of reducing the figures from the given numerical value. It defines as "to reduce a number to the desired significant figures and adjust the last reported digit".
- **Exponential Notations** is a shorthand mathematical form and written as X^n where x is multiplied itself by n time.
- **Limiting Reactant** is the reactant which is consumed entirely when the chemical reaction is completed is called Limiting reactant or Limiting reagent.
- **Reactant in excess** is the reactant which does not completely consume in the reaction and some of its amount remains unused at the end of reaction.
- **Theoretical yield** is the maximum amount of product obtained by a balanced chemical reaction.
- **Practical yield** is the amount of product formed during the experiment. This amount is usually less than theoretical yield.
- **Percent yield** is the percent ratio of practical yield and theoretical yield.

EXERCISE

Multiple Choice Questions

1. Choose the correct answer

(i) If the volume occupied by oxygen gas (O_2) at STP is 44.8dm^3, the number of molecules of O_2 in the vessels are:
(a) 3.01×10^{23} (b) 6.02×10^{23}
(c) 12.04×10^{23} (d) 24.08×10^{23}

(ii) The number of carbon atoms in 1 mole of sugar ($C_{12}H_{22}O_{11}$) are approximately:
(a) 6×10^{23} (b) 24×10^{23}
(b) (c) 60×10^{23} (d) 72×10^{23}

(iii) In the reaction $2Na + 2H_2O \rightarrow 2NaOH + H_2$, if 23g of Na reacts with excess of water, the volume of hydrogen gas (H_2) liberated at STP should be:
(a) 11.2dm^3 (b) 22.4dm^3 (c) 33.6dm^3 (d) 44.8dm^3

(iv) Which of the following sample of substances contains the same number of atoms as that of 20g calcium:
(a) 16g S (b) 20g C (c) 19g K (d) 24g Mg

(v) Which of the following statement is incorrect?
(a) The mass of 1 mole Cl_2 gas is 35.5g
(b) One mole of H_2 gas contains 6.02×10^{23} molecules of H_2
(c) Number of atoms in 23g Na and 24g Mg are equal
(d) One moles of O_2 at S.T.P occupy 22.4 dm^3 volume

(vi) For Avogadro's number, this statement is incorrect:
(a) It is the no. of particles in one moles of any substances
(b) Its numerical value is 6.02×10^{23}
(c) Its value change if temperature increases
(d) Its value change if number of moles increases

(vii) The minimum number of moles are present in:
(a) 1 dm^3 of methane gas at STP (b) 5 dm^3 of helium gas at STP
(c) 10 dm^3 of hydrogen gas at STP (d) 22.4 dm^3 of chlorine gas at STP

(viii) Number of atoms in 60g carbon are:
(a) 3.01×10^{23} (b) 3.01×10^{24}
(c) 6.02×10^{23} (d) 6.02×10^{24}

(ix) Maximum number of molecules present in the following sample of gas:
(a) 100g O_2 (b) 100g CH_4
(c) 100g CO_2 (d) 100g Cl_2

(x) Generally actual yield is:
(a) Greater than theoretical yield
(b) Less than theoretical yield
(c) Equal to the theoretical yield
(d) Some times greater and some times less than theoretical yield

Short Questions

1. Define the following:
   (i) Stoichiometry (ii) Exponential Notation (iii) Molar Volume
2. Express the following numbers in exponential notation:
   (i) 3652 (ii) 0.0231 (iii) 0.000072
3. Express the following in simple numbers.
   (i) 3.26×10^{-3} (ii) 1.921×10^{2} (iii) 1.02×10^{5}
4. Define rounding of data. Give various rules of rounding of data.

Descriptive Questions

1. Define theoretical yield, actual yield and percent yield. Why the practical yield is often less than theoretical yield?
2. What is meant by mole and Avogadro's number? How are they inter related to each other?
3. What is meant by Avogadro's number? Explain concept of mole with the help of Avogadro's number.
4. What is a Limiting reactant? How does it control the amount of product formed in a chemical reaction.

Numerical Questions

1. Calculate the number of moles and molecules in:
   (i) 38g of carbon disulphide (CS_2) (ii) 68.4g of sucrose ($C_{12}H_{22}O_{11}$)
   [Ans: (i) 0.499moles and 3.0×10^{23} molecules (ii) 0.199moles and 1.19×10^{23} molecules]
2. Ammonia gas can be produced by heating together the solid NH_4Cl and $Ca(OH)_2$.
   $$2NH_4Cl + Ca(OH)_2 \longrightarrow 2NH_3 + CaCl_2 + 2H_2O$$
   If a mixture containing 100g of each of these solids is heated, determine the limiting reactant and the mass of NH_3 gas produced.
   [Ans: Mass of NH_3 is 31.7g & Limiting reactant is NH_4Cl]
3. Aluminum chloride is used in the manufacturing of rubber. It is produced by allowing Aluminum to react with Cl_2 gas at 650°C.
   $$2Al_{(s)} + 3Cl_{2(g)} \longrightarrow 2AlCl_{3(l)}$$
   When 160g Aluminum reacts with excess of chlorine, 650g of $AlCl_3$ is produced. What is the percentage yield of $AlCl_3$?
   [Ans: 82.17%]
4. 1.6g of a sample of gas occupies a volume of 1.12 dm^3 at STP. Determine the molar mass of the substance.
   [Ans: 32g/mol]

5. Silver sulphide (Ag_2S) is an anti microbial agent. In an experiment 24.8g Ag_2S is reacted with the excess of hydrochloric acid as given in the following reaction.

$$Ag_2S_{(s)} + 2HCl_{(aq)} \longrightarrow 2AgCl_{(s)} + H_2S_{(g)}$$

Calculate the

(i) Mass of AgCl formed

(ii) Volume of H_2S produced at STP

(At. Mass of Ag is 108 g/mol and S is 32 g/mol)

[Ans: 28.7g and 2.24dm^3]

6. Calculate each of the following quantities.

(i) Number of moles in 6.4g of SO_2.

[Ans: 0.1 mol]

(ii) Mass in gram of 4.5 moles of ethyne (C_2H_2)

[Ans: 117g]

(iv) Volume in cm^3 of 38.4g O_2 gas at STP

[Ans: 26880 cm^3]

(v) Number of molecules of 126g water

[Ans: 42.14×10^{23}]

(vi) Mass in gram of 4.8×10^{24} atoms of sodium

[Ans: 183.39g]

(vi) Number of formula units in 333g of $CaCl_2$

[Ans: 18.06×10^{23}]